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Polyphosphate Water-Treatment Products: Their Effects on the Chemistry and Solubility of Lead in Potable Water Systems

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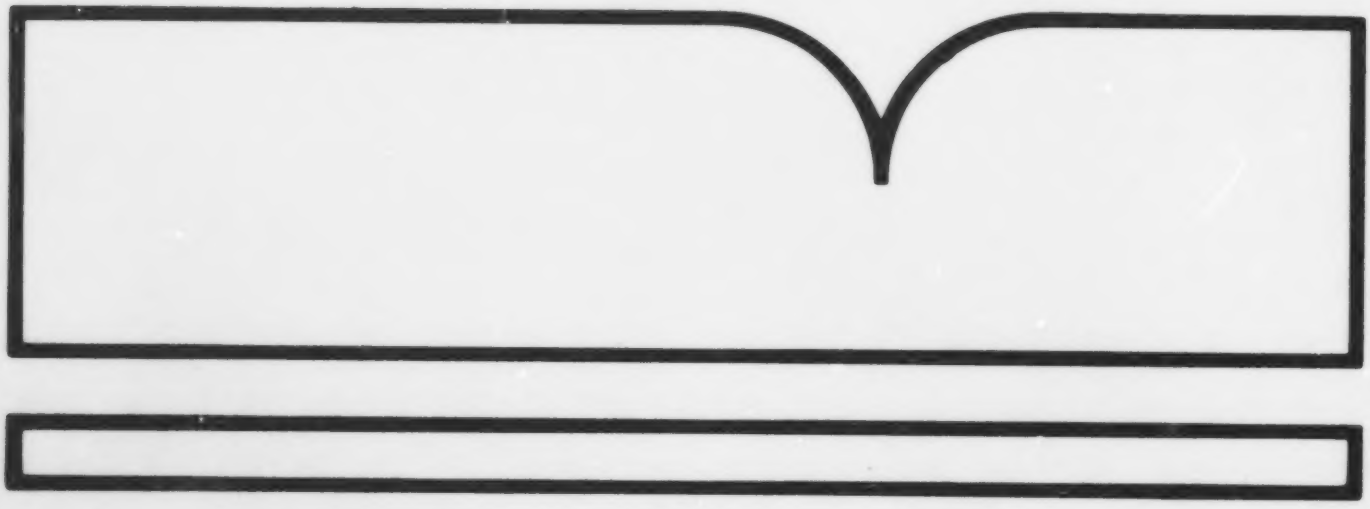

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16. ABSTRACT

Three commercial polyphosphate-containing water treatment products were evaluated for their ability to affect lead solubility in drinking water. A "Competing-Ligand Spectrophotometry" Technique was used to characterize Pb^{2+} complexation at 20 degree C in ionic strength of 0.1, pH = 8, and calcium concentration of 1×10^{-3} mol/L. The pH was buffered by HEPES, and the colorimetric reagent was 4-[2-pyridylazo]-resourcinol (PAR). Results were also compared to a DPASV method. The data was fit by a 2- ligand model. The experiments suggested a significant ability to enhance lead solubility by complexation at the pH and hardness tested.

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Table 1. MUG Reactions for E. coli Isolates Originating in Different Host Species

	Host Species		
Hours Incubation	Human	Cow	Horse
24	95.9%	94.3%	94.5%
28	99.3%	100%	100%
120	99.3%	---	---

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POLYPHOSPHATE WATER-TREATMENT PRODUCTS: THEIR EFFECTS ON THE CHEMISTRY AND SOLUBILITY OF LEAD IN POTABLE WATER SYSTEMS

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INTRODUCTION

Phosphorus-containing compounds are widely used to soften water and reduce staining due to iron oxidation ("red-water" problems). One of the main mechanisms by which these chemicals are thought to produce these effects is complexation of the metal ions Ca^{2+}, Mg^{2+}, Fe^{2+}, and Mn^{2+}. However, complexes may also be formed with toxic metal ions, including Pb^{2+}. Complexation of Pb may enhance its solubility in household plumbing systems. Complexation may therefore increase the risk of human exposure to Pb. The objective of the research described in the present paper was to characterize the interaction of Pb^{2+} with several commonly-used phosphorus-containing water-treatment products. The studies were designed to produce data that will allow modeling of the potential effect of polyphosphates on Pb solubility.

Phosphorus-containing water-treatment products can be conveniently classified as compounds containing carbon-phosphorus bonds (phosphonic acids and their salts) and those compounds that do not contain carbon. Because phosphonates are not used in treating potable waters, they are not considered in this research. The non-phosphonate compounds include the orthophosphate ion (PO_4^{3-}) and compounds produced by polymerization of orthophosphate. The simplest polyphosphate is pyrophosphate ($P_2O_7^{4-}$). Higher polymers include linear and cyclic compounds. The linear compounds can have very long chain lengths. Linear polyphosphates having up to 65 phosphate units are available commercially [Dick and Tabatabai 1987]. The general stoichiometric formula for sodium salts of linear polyphosphates having three or more phosphate units is $Na_{n+2}P_nO_{3n+1}$.

Polyphosphates form complexes with many metal ions [Manahan 1984]. This property makes these compounds useful for water treatment. For example, complexation of Fe^{2+} and Mn^{2+} reduces the tendency of water to form stains by inhibiting their oxidation. Complexation of Ca^{2+} reduces the saturation index of $CaCO_3$ and thereby inhibits scale deposition. The mechanism of scale prevention or reduction probably also involves crystal poisoning. Orthophosphate [Berner 1974] and polyphosphates [Berner and Morse 1968] inhibit the growth of crystals of calcite ($CaCO_3$).

To understand the complexation and solubility computations presented in later sections, it is helpful to review chemical equilibria. As an example of solubility equilibria, equation 1 gives the stoichiometry of

$$Pb_3(CO_3)_2(OH)_2 + 2H^+ = 3Pb^{2+} + 2CO_3^{2-} + 2H_2O \quad (1)$$

the dissolution of hydrocerrusite, a Pb-containing mineral commonly found deposited in plumbing systems [Schock 1989]. At equilibrium, the

concentrations of reactants and products are related by equation 2, where

$$K_s' = \frac{[Pb^{2+}]^3[CO_3^{2-}]^2}{[H^+]^2} \quad (2)$$

K_s' is the conditional solubility product (the true thermodynamic solubility product or equilibrium constant corrected for ionic strength). Square brackets indicate concentrations. The Pb^{2+} concentration can be computed, given pH and alkalinity, by solving equation 2, i.e. $[Pb^{2+}]$ is fixed. The saturation index (SI) of hydrocerrusite is defined as the logarithm of the ratio of the ion concentration product (actually the ion activity product) (equation 3) [Nordstrom and Munoz 1986]. At equilibrium

$$SI = 3\text{Log}[Pb^{2+}] + 2\text{Log}[CO_3^{2-}] + 2pH - \text{Log } K_s' \quad (3)$$

or saturation, the value of SI is zero. A positive value of SI indicates oversaturation, while a negative value indicates undersaturation.

As an example of a complexation reaction, equation 4 shows the

$$Pb^{2+} + L = PbL \quad (4)$$

stoichiometry of the reaction of Pb^{2+} with a ligand L to form the complex PbL. At equilibrium, the concentrations of reactants and products are related by equation 5, where K' is the conditional stability constant.

$$K' = \frac{[PbL]}{[Pb^{2+}][L]} \quad (5)$$

(The true thermodynamic stability constant corrected for ionic strength.) The size of the stability constant indicates the strength of the complex.

For a hypothetical solution containing only one ligand that complexes Pb^{2+} (real waters contain several Pb-complexing ligands), the mass balance for dissolved Pb is given by equation 6, where Pb_{tot} is the total dissolved

$$Pb_{tot} = Pb^{2+} + PbL \quad (6)$$

concentration of Pb. From equations 2 and 6, increasing the total concentration of L or the value of K', decreases the free Pb^{2+} concentration. For a fixed value of Pb^{2+} (e.g. in equilibrium with hydrocerrusite) increasing the concentration of L or the value of K', increases Pb_{tot}.

Metal-ligand interactions depend on many physical and chemical factors. Polyphosphates form complexes with many metal ions. Therefore, the complexation of one metal ion, such as Pb^{2+}, depends on competition with other metals (Ca^{2+}, Mg^{2+}, Fe^{2+}, Mn^{2+}) for available ligands. Polyphosphates are weak acids. Therefore, polyphosphate metal binding is a function of pH. All ionic equilibria, including polyphosphate reactions, are affected by the ionic strength. Most equilibrium constants are functions of temperature.

The objective of the present research was to characterize Pb^{2+} complexation by representative polyphosphate water-treatment chemicals under one set of chemical and physical conditions that are representative of some potable waters. The Pb^{2+} complexation was modeled by assuming the diluted polyphosphate products contained two ligands that formed 1:1 complexes with Pb^{2+} (equation 5). The polyphosphate products may actually have contained more than two Pb-complexing ligands. However, the purpose of

the model was to describe Pb complexation in terms of a simple model. The "ligand concentrations" and "stability constants" produced by fitting the model to the experimental data are not concentrations of actual species or true thermodynamic stability constants. They are parameters that can be used as ligand concentrations and stability constants in chemical-equilibrium calculations. Their use is only valid under the conditions (pH, temperature, ...) under which they were determined.

MATERIALS AND METHODS

The water treatment products studied were chosen to be representative of the many products available for water treatment. CalciQuest (Waltron Ltd., Whitehouse, NJ) and Aqua-Mag (Kjell Water Consultants, Inc., Beloit, WI) are proprietary solutions of sodium polyphosphates. C-5 (Calgon Corp., Pittsburgh, PA) is a solution containing sodium ortho-, pyro-, and tripolyphosphates. C-39 and TG-10 (Calgon Corp.) are sodium/zinc polyphosphates in solid and solution form, respectively. All other chemicals were reagent grade.

The ionic strength was kept constant at 0.1 eq L^{-1} by adding $NaNO_3$. This ionic strength is higher than for most potable waters. However, it allowed constant experimental conditions. Also, published stability constants, which were needed to analyze experimental data, were available for ionic strength 0.1, but were not always available for lower ionic strengths. All experiments were performed at room temperature (20 C). The pH was buffered using HEPES (N-[2-hydroxyethyl]piperazine-N'-2-ethanesulfonic acid), which has a negligible affinity for metal ions [Good et al. 1966]. All solutions were made CO_2-free to avoid carbonate complexation and precipitation of Pb carbonates. For all experiments, the solution being titrated contained $1x10^{-3}$ M Ca (hardness 100 mg L^{-1} as $CaCO_3$). The Ca was used to simulate competition between Pb^{2+} and major metal ions in potable waters. Lead contamination from reagents was reduced below detectable levels by ion exchange using Chelex-100 (Bio-Rad, Richmond, CA) or MnO_2 adsorption [Van den Berg 1982].

Complexation of Pb^{2+} was studied by complexometric titrations monitored by two independent analytical methods: differential pulse anodic stripping voltammetry (DPASV) and competing-ligand spectrophotometry (CLS). The CLS experiments were batch titrations. The colorimetric reagent was 4-[2-pyridylazo]-resorcinol (PAR). The PAR concentration was $1x10^{-4}$ M for experiments using Aqua-Mag and C-5 and $2.5x10^{-4}$M for experiments using CalciQuest. PAR forms strong, highly colored complexes with Pb^{2+} and many other metals, but not with Ca^{2+}. The Pb^{2+} complexes have maximum absorbances at 510 nm, while PAR has its maximum absorbance at 410 nm [Pollard et al. 1959]. The extinction coefficients of PAR and the Pb-PAR complexes were determined before titrations of any water-treatment products were performed. Because Zn^{2+} also forms colored complexes with PAR, CLS was not used to study Pb^{2+} complexation by the Zn-phosphate products C-39 and TG-10.

The DPASV experiments were performed using a BAS-100 potentiostat (Bioanalytical Systems, East Lafayette, IN) and an EG&G-Princeton Applied Research (Princeton, NJ) model 303A static mercury drop electrode with a model 305 electronic stirrer. Teflon electrochemical cells were used to avoid sorptive losses of Pb. The experimental protocol and method of data analysis of Van den Berg [1984] were used in the DPASV experiments. The total Pb concentration was kept constant at $1.0x10^{-7}$ M to avoid precipitation of $Pb(OH)_2$. The concentration of water-treatment chemical was varied. Thus, the Pb^{2+} was titrated with the polyphosphate product. For CalciQuest and AquaMag the dilution factor ranged from $1x10^{-7}$ to

$1x10^{-3}$. The manufacturers' recommended doses for both products were included in this range. For C-39 and TG-10 the concentration ranged from 0.01 to 100.0 mg product per liter.

ASV gives a peak-shaped output. In systems involving Pb complexation by a single ligand, the potential of the peak maximum is given by equation 7.

$$E_p = E_{p,0} + C - RT\ \mathrm{Ln10}(1 + K[L])/2F \qquad (7)$$

where E_p is the peak potential and $E_{p,0}$ is the peak potential for no significant complexation (i.e. very low L concentration). Ln10 is the natural logarithm of 10. C is a collection of constants (diffusion coefficients, etc.). The other symbols have their usual meanings. The value of RT Ln10/2F is approximately 29 mV at 20°C. For very low ligand concentrations, the value of E_p is constant. For high ligand concentrations, E_p decreases linearly with a slope of -29 mV (log [L])$^{-1}$. There is a transitional region whose shape and extent depend on K. DPASV results were analyzed using equation 8. Note that the peak potential depends on the product of the stability constant and the ligand concentration. The two parameters cannot be determined independently by DPASV. Therefore, DPASV is most useful in qualitatively confirming that polyphosphates complex Pb^{2+} and in possibly testing the complexation parameters obtained by CLS.

The CLS complexometric titrations were analyzed using discrete one-and two-ligand models. These models generate parameters that can be used as stability constants and ligand concentrations in chemical-equilibrium calculations to describe the degree of Pb^{2+} and Ca^{2+} binding. This modeling approach has been applied to the complexation of metal ions by natural aquatic organic matter [Cabaniss and Shuman 1988]. In comparisons with electrostatic, multidentate, affinity-spectrum, and normal-distribution models, the multiple-discrete-ligand model gives superior fits to experimental data. Discrete-ligand models produce parameters that can be easily used in existing chemical-equilibrium computer programs [Turner et al. 1986].

The model for CLS experiments is given in equation 8, where A and ϵ stand

$$A_{410} = \epsilon_{410,HPAR}[HPAR] + \epsilon_{410,PbHPAR}[PbHPAR] + \epsilon_{410,Pb(HPAR)_2}[Pb(HPAR)_2] \qquad (8a)$$

$$A_{510} = \epsilon_{510,HPAR}[HPAR] + \epsilon_{510,PbHPAR}[PbHPAR] + \epsilon_{510,Pb(HPAR)_2}[Pb(HPAR)_2] \qquad (8b)$$

for absorbance and extinction coefficient, respectively. Numerical subscripts indicate wavelengths. Other subscripts (e.g. HPAR) indicate species. As examples, A_{410} is the absorbance at 410 nm and $\epsilon_{410,HPAR}$ is the extinction coefficient of HPAR at 410 nm. The extinction coefficients of HPAR$^-$ were determined by measuring the absorbances of PAR solutions. (The predominant chemical form of PAR at the pH of the CLS experiments is HPAR$^-$.) The extinction coefficients of PbHPAR$^+$ and $Pb(HPAR)_2^0$ were determined by measuring the absorbances of solutions containing only Pb^{2+} and PAR. The concentrations of HPAR$^-$, PbHPAR$^+$, and $Pb(HPAR)_2$ were computed using stability constants from the critical compilation of Martell and Smith [1974a]. Multilinear regression was used to compute the extinction coefficients. The pH and ionic strength of the calibration were the same as for the titrations of polyphosphates.

Fitting of the models to the CLS data was performed using simplex optimization [Caceci and Cacheris 1982] to minimize the sum of squares of

residuals (equation 9), where "comp" in the subscripts indicates computed

$$\chi^2 = \Sigma((A_{410} - A_{410,comp})^2 + (A_{510} - A_{510,comp})^2) \qquad (9)$$

absorbances. The absorbances were computed according to equation 8. The species concentrations for substitution into equation 9 were computed using a method based on Newton-Raphson iteration [Morel and Morgan 1972, Taylor et al. 1988]. The chemical equilibrium problem was solved for each data point in each iteration. The standard deviations of the complexation parameters were estimated by the method of Phillips and Eyring [1988]. The DPASV data was modeled using the nonlinear least squares regression program MINSQ (Micro-Math Scientific Software, Salt Lake City, UT).

The uncertainties in total dissolved Pb concentrations computed using the complexation parameters were estimated using Monte-Carlo simulation. The familiar first-order approximation for error propagation [Roddick 1987] could not be used because the computed total Pb concentration was not a linear function of the parameters. The method used is outlined as follows. Assume the errors in the complexation parameters are normally distributed. Let p be the vector of complexation parameters and V be the variance-covariance matrix. (For each model "ligand" the concentration and stability constant were highly correlated.) Let LL^T be the Cholesky decomposition of V. Let n be a vector of normally distributed random numbers with means of zero and unit variances. Then the perturbation vector Δp was computed according to equation 10. The elements of

$$\Delta p = L\,n \qquad (10)$$

the perturbed vector of parameters, $p + \Delta p$, were used to compute the total dissolved Pb. This process was repeated 1000 times and the standard deviation of computed total dissolved Pb concentrations was computed.

The CLS method was tested by titrating nitrilotriacetic acid (NTA). The stability constant of the PbNTA$^-$ complex is well known [Martell and Smith 1974a]. The concentration of NTA was $1x10^{-6}$ M, which gave a complexing capacity similar to those measured in the titrations of the polyphosphates. The test measured the ability of the CLS method to determine both ligand concentration and stability constant.

RESULTS AND DISCUSSION

The results of the CLS method validation are listed in Table 1. Both the ligand concentration and the stability constant compared very well with the true values. The results were reproducible. Therefore, the CLS method produced valid results.

The results of a CLS experiment involving Aqua-Mag are shown in Figure 1. The dilution factor of $1x10^{-6}$ was in the range typically used in treating drinking water. The PAR concentration was $1x10^{-6}$ M. The absorbance of the Aqua-Mag-PAR-Pb solution at 510 nm was less than the absorbance of the PAR-Pb solution. The polyphosphate ligands formed complexes with Pb^{2+} ions in competition with HPAR$^-$. As a result, the concentrations of PbHPAR$^+$ and $Pb(HPAR)_2^0$, which absorb at 510 nm, were reduced. Similarly, the absorbance of the Aqua-Mag-PAR-Pb solution at 410 nm was greater than the absorbance of the PAR-Pb solution. Aqua-Mag reduced the amount of Pb-PAR complex formation and increased the HPAR$^-$ concentration. There was also evidence of competition for Pb^{2+} between CalciQuest and C-5 and the strong complexing reagent PAR.

The Pb^{2+} complexation parameters determined by fitting two-ligand models to the CLS titrations are presented in Table 2. The results were reproducible as shown by the complexation parameters for duplicate titrations. Some of the parameters had rather large standard deviations. However, the stability constants and ligand concentrations were often highly correlated ($r \sim -1$). Therefore the standard deviation of the Pb^{2+} concentrations computed using the parameters are much smaller than may be expected based solely on the standard deviations.

The results of the DPASV experiments had many of the features predicted by equation 7. At low ligand concentrations (high dilution of product) the peak potential was constant. At high ligand concentrations the peak potential decreased linearly with the logarithm of ligand concentration. However, the fit was (visually) poor. The slopes of of the linear portions of all of the titration curves but one were significantly different (95% confidence level) from -29 mV (log dilution factor)$^{-1}$. The DPASV experiments did qualitatively confirm that Aqua-Mag, CalciQuest, C-5, C-39, and TG-10 all formed complexes with Pb^{2+}.

POTENTIAL EFFECTS OF POLYPHOSPHATE COMPLEXATION ON LEAD SOLUBILITY

All polyphosphate water-treatment products studied formed strong complexes with Pb. The complexes were formed under conditions typical of many potable waters, i.e. pH 8.0 and $1x10^{-3}$ M Ca, hardness 100 mg L^{-1} as $CaCO_3$). Therefore, polyphosphate complexation may be an important factor in speciation of Pb in domestic plumbing systems.

Chemical equilibrium computations were performed to estimate the effects of polyphosphate Pb complexation on Pb solubility. The free Pb^{2+} concentration in equilibrium with hydrocerussite, $Pb_3CO_3(OH)_2$, and hydroxypyromorphite, $Pb_5(PO_4)_3OH$. Polyphosphates hydrolyze (revert) to orthophosphate in aqueous solution. Thus, a Pb phosphate may precipitate rather than hydrocerrusite. Orthophosphate concentrations equivalent to up to 20% reversion of the polyphosphates were considered. The minimum free Pb^{2+} (in equilibrium with hydrocerrusite or hydroxypyromorphite) was used to compute Pb solubility. The concentrations of the complexes of Pb^{2+} formed with CO_3^{2-}, OH^-, and polyphosphates were computed and added together to give the total soluble Pb concentration. Alkalinity values of 5, 50, and 150 mg L^{-1} were used to simulate soft and moderately hard waters. (Scale formation is, of course, not a serious problem with soft waters. However, polyphosphate products are frequently used to control "red water" and corrosion in soft-water areas. Thus, the low alkalinity value was included in the Pb solubility computations.) The pH and ionic strength for the computations were the same as in the complexation experiments.

The results of the solubility computations are shown in Figure 2. Aqua-Mag, CalciQuest, and C-5 caused increases in total Pb concentrations at all three alkalinity values through the formation of Pb complexes. Aqua-Mag at a dilution of $1x10^{-5}$ increased the total soluble Pb concentration by factors of 1.4 to 2.9 compared to polyphosphate-free solutions for all three alkalinity values. For CalciQuest diluted by a factor of $2.5x10^{-6}$, the total Pb concentrations were somewhat less than for Aqua-Mag. The observed increases in total soluble Pb produced by C-5 at a dilution of $1x10^{-5}$ were greater than for Aqua-Mag or CalciQuest. However, the C-5 dilution may have been less than typically used to treat drinking water.

Complexation of Pb^{2+} by polyphosphates may influence both the kinetics and equilibria of Pb precipitation-dissolution reactions by greatly reducing the fraction of dissolved Pb in the free Pb^{2+} form. Crystal poisoning by

polyphosphates may inhibit the precipitation of Pb carbonates as it inhibits the precipitation of $CaCO_3$. The rates of precipitation (in the absence of crystal poisoning) and dissolution are functions of the saturation index. Adding polyphosphates to water would, for a given total dissolved Pb concentration, decrease the saturation index of hydrocerrusite relative to the value before the polyphosphate addition. Decreasing the saturation index would decrease the rate of precipitation from an oversaturated solution or increase the rate of dissolution into an undersaturated solution. At equilibrium, the total dissolved concentration of Pb in water containing polyphosphates would be higher than in the same water (i.e. the same pH, alkalinity, ...) without polyphosphates.

The use of polyphosphate-based products for treating potable water may increase the risk of human exposure to Pb through drinking water under certain conditions. Consider a plumbing system in which Pb^{2+} is introduced into solution by corrosion and is removed from solution by precipitation of hydrocerrusite. Adding polyphosphates to the water would increase the dissolved Pb concentration by raising the saturation concentration or by poisoning the nucleation sites on the existing hydrocerrusite deposits. The degree of increase in the dissolved Pb concentration would depend on the corrosion rate, the pH and alkalinity of the water, and the concentration and Pb complexing properties of the polyphosphate. However, polyphosphates may inhibit corrosion. Therefore, the increase in dissolved Pb due to inhibition of precipitation would last only until the corrosion inhibition became effective. After cessation of corrosion, Pb would be introduced into solution mainly by dissolution of the hydrocerrusite deposits. The rate of dissolution may be affected by complexation of Pb^{2+} by the polyphosphates.

SUGGESTIONS FOR FURTHER RESEARCH

The results of the present research suggest that complexation of Pb by polyphosphates may significantly affect the chemistry and solubility of Pb in domestic plumbing systems. The complexation parameters determined in the present research are useful for modeling Pb chemistry under conditions that are representative of some potable waters. However, much more research is needed to assess the potential impact of polyphosphates on Pb solubility in all plumbing systems.

The results of the present research should be confirmed by repeating the measurements made with the CLS method and by independent methods. Methods for studying Pb complexation by zinc-containing water treatment products should be developed, possibly based on ion exchange equilibria. The Pb-complexing properties of many more water treatment products should be characterized. The effects of pH, ionic strength, and competing ions on Pb complexation should be determined. Complexation of other metals (Cu^{2+}, Cd^{2+}, Zn^{2+}) by polyphosphates should be studied. The effects of polyphosphates on the precipitation and dissolution kinetics of hydrocerrusite and other Pb minerals should be determined.

Competing-ligand spectrophotometry using PAR is a method that can be used by virtually any laboratory for studying Pb chemistry. The only instruments needed are a spectrophotometer and a pH meter. As a result, it may be possible to perform site-specific testing of water-treatment chemicals for Pb complexation. However, special precautions must be taken to avoid contamination of reagents and glassware by Pb and other metals. Also, much more work before is needed to verify the method before it can be used routinely.

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TABLE 1. COMPLEXATION PARAMETERS DETERMINED BY COMPETING-LIGAND SPECTROPHOTOMETRY FOR A TEST CASE: 10 μM NTA

	LOG CONCENTRATION*	LOG STABILITY CONSTANT*
TRUE OR ACCEPTED VALUE**	-5.00	11.34
EXPERIMENTAL VALUES	-5.03 (0.01)	11.85 (0.90)
	-5.01 (0.01)	11.51 (0.14)

* STANDARD DEVIATIONS IN PARENTHESES
** TRUE VALUE OF CONCENTRATION. ACCEPTED VALUE OF STABILITY CONSTANT [MARTELL AND SMITH 1974a]

TABLE 2. PARAMETERS FOR FORMATION OF LEAD COMPLEXES FOR NON-ZINC-CONTAINING POLYPHOSPHATE WATER TREATMENT PRODUCTS DETERMINED BY COMPETING-LIGAN SPECTROPHOTOMETRY

PRODUCT	DILUTION ($X10^4$)	STABILITY CONSTANTS* LOG K_1	LOG K_2	LIGAND CONCENTRATIONS* LOG $L_{1,TOT}$	LOG $L_{2,TOT}$
AQUA-MAG	1.0	9.97 (0.59)	3.97 (0.88)	-6.41 (0.16)	-3.09 (0.87)
	3.0	10.52 (0.14)	6.45 (0.11)	-5.73 (0.04)	-4.59 (0.06)
CALCIQUEST	0.25	9.62 (0.18)	5.38 (0.28)	-6.07 (0.06)	-4.78 (0.23)
	1.0	9.98 (0.38)	5.40 (0.09)	-5.98 (0.10)	-4.38 (0.11)
	1.0	9.95 (0.17)	5.00 (0.22)	-6.08 (0.06)	-3.99 (0.23)
	3.0	10.68 (0.09)	7.85 (0.15)	-5.30 (0.04)	-4.93 (0.04)
C-5	1.0	10.43 (0.07)	6.44 (0.27)	-5.48 (0.02)	-4.49 (0.20)
	1.0	10.04 (0.34)	7.35 (0.20)	-5.50 (0.13)	-4.94 (0.04)

* STANDARD DEVIATIONS IN PARENTHESES

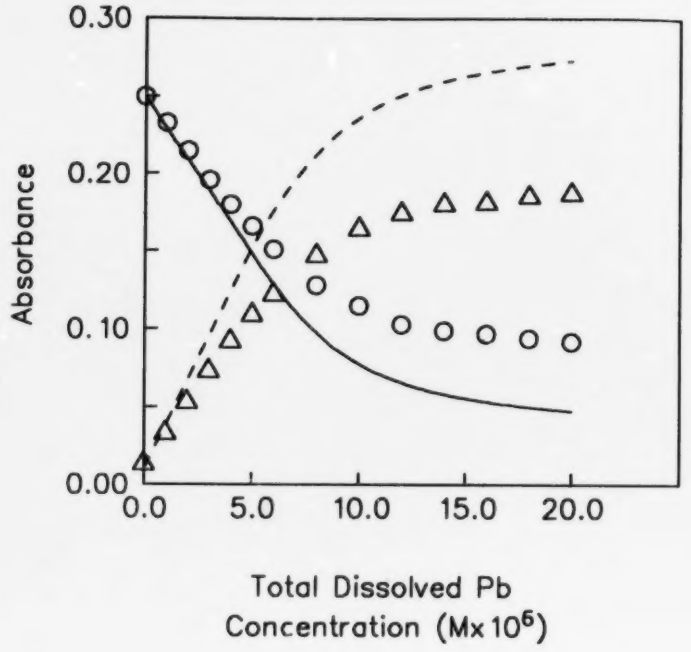

Figure 1. Competing-ligand spectrophotometric titration of Aqua Mag, $1x10^{-5}$ dilution. Total PAR $1x10^{-5}$ M.
— No Aqua-Mag, 410 nm. o With Aqua-Mag, 410 nm.
-- No Aqua-Mag, 510 nm. Δ With Aqua-Mag, 510 nm.

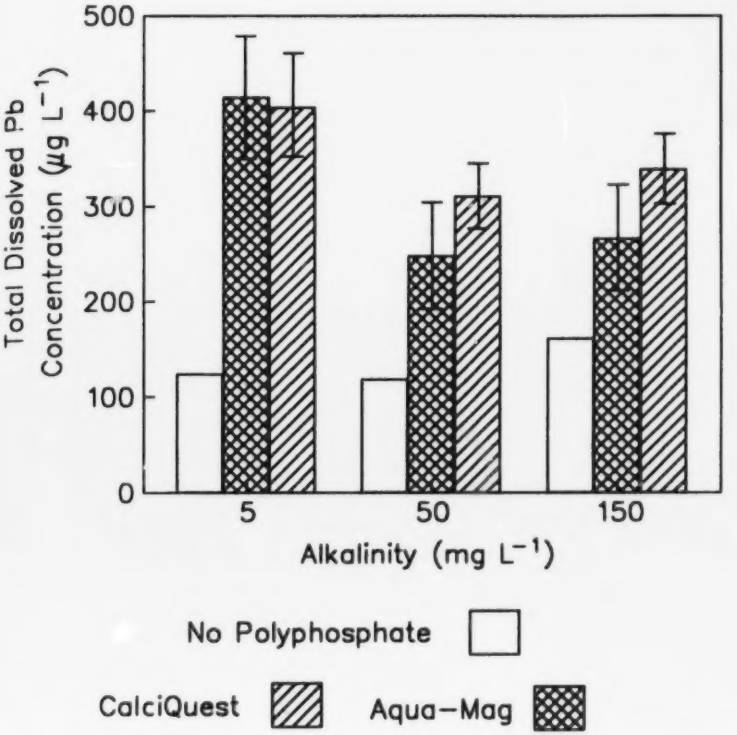

Figure 2. Effects of polyphosphate complexation on computed Pb solubility. Product dilutions: Aqua-Mag $1x10^{-5}$, CalciQuest $2.5x10^{-6}$. Ionic strength 0.1 eq L^{-1}, pH 8.0.